

Homework #1

1. A solution of sodium bicarbonate is prepared by adding 45.00 g of sodium bicarbonate to a 1.00 L volumetric flask and adding distilled water until it reaches the 1.00 L mark. What is the concentration of sodium bicarbonate in units of (a) milligrams per liter, (b) molarity, (c) normality, and (d) milligrams per liter as CaCO₃?

Given: 45.000 g of sodium bicarbonate in 1.00 L of water.

Solution:

1 point

- a. The resulting concentration of the solution is 45 g/L. Converting this to mg/L

$$(45 \text{ g/L})(1000 \text{ mg/g}) = 45,000 \text{ mg/L} = 4.5 \times 10^4 \text{ mg/L}$$

- b. To find the molarity, the molecular weight of NaHCO₃ must be found

$$\begin{aligned} \text{Na} &= 22.99 \times 1 = 22.99 \\ \text{H} &= 1.008 \times 1 = 1.008 \\ \text{C} &= 12.01 \times 1 = 12.01 \\ 3\text{O} &= 16.00 \times 3 = 48.00 \\ \Sigma &= 84.01 \text{ g/mol} \end{aligned}$$

Then, the molarity

$$(45 \text{ g/L})(1 \text{ mol}/84 \text{ g}) = 0.536 \text{ M} = 0.54 \text{ M}$$

- c. The equivalent weight of NaHCO₃ is its GMW divided by the number of hydrogen ions transferred. In this case $z = 1$ because Na⁺ is replaced by 1 H. Thus normality (N) is $z \cdot M$, and in this case $z = 1$

$$(0.536 \text{ M})(1 \text{ N/M}) = 0.536 \text{ M} = 0.54 \text{ N}$$

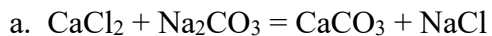
- d. Change to units of mg/L as CaCO₃, using conversion (eq. 2-87 and lecture 4, slide 7).

$$45,000 \text{ mg/L} * \left(\frac{50 \text{ mg/L as CaCO}_3}{84 \text{ mg/L as NaCO}_3} \right) = 26,785 \text{ mg/L} = 2.7 \times 10^4 \text{ mg/L as CaCO}_3$$

2. the following 5 chemical equations:

Solution:

1 point

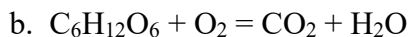
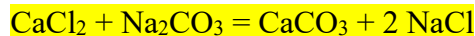


<u>Elements</u>	<u>Reactants</u>	<u>Products</u>
Ca	1	1
Cl	2	1
Na	2	1
C	1	1
O	3	3

Note that we are short 1 Na, so multiply product NaCl by 2

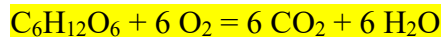
<u>Elements</u>	<u>Reactants</u>	<u>Products</u>
Ca	1	1
Cl	2	2
Na	2	2
C	1	1
O	3	3

This yields the balanced equation:

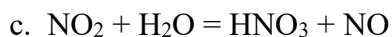


<u>Elements</u>	<u>Reactants</u>	<u>Products</u>
C	6	1
H	12	2
O	8	3

Note that we are short 5 C, 10 H and 5 O. Multiply product CO_2 by 6 and H_2O by 6. This balances C and H on each side, and leaves a difference of 10 O. Multiply the reactant O_2 to balance equation:

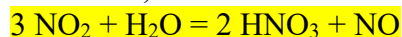


<u>Elements</u>	<u>Reactants</u>	<u>Products</u>
C	6	6
H	12	12
O	18	18

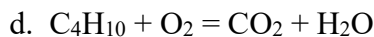


<u>Elements</u>	<u>Reactants</u>	<u>Products</u>
N	1	2
H	2	1
O	3	4

Note that H is short 1, therefore multiply the product HNO₃ by 2. This causes N to be out of balance, and the reactant NO₂ should be multiplied by 3.

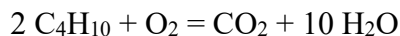


<u>Elements</u>	<u>Reactants</u>	<u>Products</u>
N	3	3
H	2	2
O	7	7



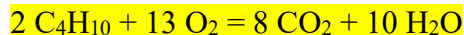
<u>Elements</u>	<u>Reactants</u>	<u>Products</u>
C	4	1
H	10	2
O	2	3

Note that H is out of balance, but in order to keep the number of O even, we must multiply the product H₂O by 10. As a result, the reactant C₄H₁₀ must be multiplied by 2.

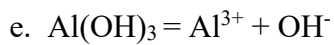


<u>Elements</u>	<u>Reactants</u>	<u>Products</u>
C	8	1
H	20	20
O	2	12

However, the reaction is still not balanced. CO₂ must be multiplied by 8 to maintain an equal number of C. As a result, O is left unbalanced. Multiply O₂ by 13.



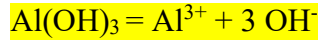
<u>Elements</u>	<u>Reactants</u>	<u>Products</u>
C	8	8
H	20	20
O	26	26



<u>Elements</u>	<u>Reactants</u>	<u>Products</u>
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Al	1	1
O	3	1
H	3	1

Note that Al is in balance, however we are missing two OH groups. Multiply the product OH⁻ by 3.



<u>Elements</u>	<u>Reactants</u>	<u>Products</u>
Al	1	1
O	3	3
H	3	3

3. You've just prepared a solution by dissolving 20 mg sodium sulfide (Na₂S), and 30 mg potassium sulfate dihydrate (K₂SO₄•2H₂O) in 1 Liter of distilled water.

1 point

- What is the molar concentration of sodium sulfide in this solution?
- What is the equivalent concentration of sodium sulfide in this solution?
- What is the molar concentration of potassium in this solution?
- What is the concentration of total sulfur in this solution in mg/L?
- What is the concentration of reduced sulfur (i.e., S(-II)) in this solution in mg/L?
- What is the theoretical TDS of this solution in mg/L?

Answers:

Assemble essential data:

	<u>Atomic</u> <u>Wt</u>
Na	22.99
S	32.06
K	39.1
O	16
H	1.01

Calculate molar quantities and percentages

GFW	Amount Added mg	mMoles	%water	%Sulfur	%S(-II)
78.04 Na ₂ S	20	0.256279	0.00%	41.08%	41.08%
210.3 K ₂ SO ₄ 2H ₂ O	30	0.142653	17.14%	15.24%	0.00%

a. Molar Concentration

$$GFW = 2 * 22.99 + 32.06 = 78.04$$

$$Molarity = \frac{mass/L}{GFW} = \frac{0.020g/L}{78.04g/mole} = 0.000256moles/L$$

or
0.26mM

b. Equivalent Concentration

Since z=2 for the compound (i.e., there are 2 positive charges Na⁺ and Na⁺, and 2 negative charges S⁻²), then GEW = GFW/2 = 39.02, This also makes the equivalent concentration twice the molar concentration, so:

$$Normality = \frac{mass/L}{GEW} = \frac{0.020g/L}{39.02g/mole} = 0.000513N$$

or
0.51 meq/L

c. Molar Concentration of K

first determine the molar concentration of the potassium sulfate

$$GFW = 2 * 39.1 + 32.06 + 4 * 16 + 2 * (16 + 2 * 1.10) = 210.3$$

$$Molarity = \frac{mass/L}{GFW} = \frac{0.030g/L}{210.3g/mole} = 0.000143moles/L$$

or
0.143mM

Then translate this to a concentration of potassium alone

$$Molarity = \frac{0.000143moles - K_2SO_4 \bullet 2H_2O}{L} \left(\frac{2moles - K}{1mole - K_2SO_4 \bullet 2H_2O} \right) = 0.000285moles - K / L$$

or
0.29mM

d. Concentration of Total Sulfur

One approach is to determine total molar sulfur concentration and then convert to mg/L units.

$$\begin{aligned} \text{Molarity} &= \frac{0.000256 \text{ moles} - \text{Na}_2\text{S}}{L} \left(\frac{1 \text{ mole} - S}{1 \text{ mole} - \text{Na}_2\text{S}} \right) + \frac{0.000143 \text{ moles} - \text{K}_2\text{SO}_4 \bullet 2\text{H}_2\text{O}}{L} \left(\frac{1 \text{ mole} - S}{1 \text{ mole} - \text{K}_2\text{SO}_4 \bullet 2\text{H}_2\text{O}} \right) \\ &= 0.000399 \text{ moles} - S / L \\ \text{Concentration} &= \frac{0.000399 \text{ mole} - S}{L} \left(\frac{32.06 \text{ g} - S}{\text{mole} - S} \right) = 0.0128 \text{ g} / L \\ &= 12.8 \text{ mg/L} \end{aligned}$$

e. Concentration of Reduced Sulfur

The reduced sulfur is only the S in Na₂S. This is sulfur in its (-II) oxidation state.

$$\begin{aligned} \text{Molarity} &= \frac{0.000256 \text{ moles} - \text{Na}_2\text{S}}{L} \left(\frac{1 \text{ mole} - S}{1 \text{ mole} - \text{Na}_2\text{S}} \right) \\ &= 0.000256 \text{ moles} - S / L \\ \text{Concentration} &= \frac{0.000256 \text{ mole} - S}{L} \left(\frac{32.06 \text{ g} - S}{\text{mole} - S} \right) = 0.0082 \text{ g} / L \\ &= 8.2 \text{ mg/L} \end{aligned}$$

f. What is the TDS

Determine the total mass of salts added, minus any water (i.e., water of crystallization).

$$\begin{aligned} \text{TDS} &= (20 \text{ mg} - \text{Na}_2\text{S}) \left(1 - \frac{0 \text{ g} - \text{H}_2\text{O}}{78.04 \text{ g} - \text{Na}_2\text{S}} \right) + (30 \text{ mg} - \text{K}_2\text{SO}_4 \bullet 2\text{H}_2\text{O}) \left(1 - \frac{36.04 \text{ g} - \text{H}_2\text{O}}{210.3 \text{ g} - \text{K}_2\text{SO}_4 \bullet 2\text{H}_2\text{O}} \right) \\ &= 44.9 \text{ mg/L} \end{aligned}$$

4. If 200 mg of HCl is added to water to achieve a final volume of 1.00 L, what is the final pH?

pH of HCl solution

Given: 200 mg of HCl in 1.00 L

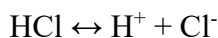
Solution:

- a. Calculate molarity

GMW of HCl = 36.45

$$\left(\frac{200 \text{ mg}}{\text{L}}\right) \left(10^{-3} \frac{\text{g}}{\text{mg}}\right) \left(\frac{1}{36.45 \text{ g/mol}}\right) = 5.487 \times 10^{-3} \text{ M of HCl}$$

- b. Moles of H⁺ on ionization



so 1 mole HCl = 1 mole H⁺

$$[\text{H}^+] = 5.476 \times 10^{-3} \text{ M}$$

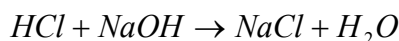
- c. Using Eqn 2-38

$$\text{pH} = -\log(5.487 \times 10^{-3}) = 2.26$$

5. What amount (mass, in mg) of NaOH (a strong base), would be required to neutralize the acid in previous problem?

From the problem above, we now know that 200 mg/L of HCl is $5.476 \times 10^{-3} \text{ M}$

Since HCl has a single hydrogen ion it donates to water, and NaOH has a single hydroxide that can neutralize a single hydrogen ion, the stoichiometry of neutralization is 1:1



Therefore we need $5.476 \times 10^{-3} \text{ M}$ of NaOH

The GMW of NaOH is $22.99 + 16 + 1.008 = 40 \text{ g/M}$

0.5
points

0.5
points

So the amount of NaOH is $40 \times 5.476 \times 10^{-3} \text{ M} = 0.219 \text{ g}$

6. The concentration of a chemical degrades in water according to first-order kinetics. The degradation constant is 0.2 day^{-1} . If the initial concentration is 100.0 mg/L , how many days are required for the concentration to reach 0.14 mg/L ? Also calculate the half-life ($t_{1/2}$) for this decay reaction.

Given: 1st order kinetics, rate of 0.2 d^{-1} , initial concentration = 100 mg/L .

Solution:

- a. Time to reach 0.14 mg/L by 1st order kinetics

$$\ln \frac{C}{C_0} = -kt$$

$$\ln \frac{0.14}{100} = -(0.2)t$$

$$-6.57 = -0.2t$$

$$t = 32.85 \text{ or } 32.9 \text{ d}$$

- b. Half Life

$$t_{1/2} = \frac{\ln 2}{K} = \frac{0.693}{0.2 \text{ d}^{-1}} = 3.5 \text{ d}$$

7. Each mole of $\text{CaF}_2(\text{s})$ dissolved yields 1 mole of Ca^{2+} and 2 moles of F^- (fluoride). The solubility product of calcium fluoride (CaF_2) is 3×10^{-11} at 25°C . Could a fluoride concentration of 1.0 mg/L be obtained in water that contains 200 mg/L of calcium? Show your work.

Given: Solubility product of $\text{CaF}_2 = 3 \times 10^{-11}$, $\text{F} = 1.0 \text{ mg/L}$ and $\text{Ca} = 200 \text{ mg/L}$

Solution:

- a. Convert Ca and F to moles/L

$$\text{Ca} = \frac{200 \text{ mg/L}}{(40.08 \text{ g/mole})(1000 \text{ mg/g})} = 4.99 \times 10^{-3} \text{ moles/L}$$

1

point

not
graded

$$F = \frac{1.0 \text{ mg/L}}{(18.998 \text{ g/mole})(1000 \text{ mg/g})} = 5.26 \times 10^{-5} \text{ moles/L}$$

b. Calculate solubility of F with 200 mg/L of Ca in solution.

$$K_s = [\text{Ca}][\text{F}]^2$$

$$K_s = [4.99 \times 10^{-3}][\text{F}]^2 = 3.00 \times 10^{-11}$$

$$[\text{F}] = \left(\frac{3.00 \times 10^{-11}}{4.99 \times 10^{-3}}\right)^{1/2} = 7.75 \times 10^{-5} \text{ moles/L}$$

c. Since 7.75×10^{-5} is greater than 5.26×10^{-5} , the 1.0 mg/L of F will be soluble.